

CHM129  
Entropy and Gibbs Free Energy

1. Discuss with your neighbors whether the signs of  $\Delta S$  for the following processes are expected to be positive or negative:

a) ice melts

$$\Delta S > 0$$

b) the temperature of a solid is lowered by 25°C.

$$\Delta S < 0$$

c) water evaporates from a beaker

$$\Delta S > 0$$

d)  $\text{HCl}_{(g)}$  dissolves in water

$$\Delta S < 0$$

e)  $4 \text{NH}_{3(g)} + 5 \text{O}_{2(g)} \rightarrow 4 \text{NO}_{(g)} + 6 \text{H}_2\text{O}_{(g)}$

$$\Delta S > 0$$

f)  $\text{CaO}_{(s)} + \text{CO}_{2(g)} \rightarrow \text{CaCO}_{3(s)}$

$$\Delta S < 0$$

g)  $\text{CH}_3\text{OH}_{(g)} \rightarrow \text{CH}_3\text{OH}_{(l)}$

$$\Delta S < 0$$

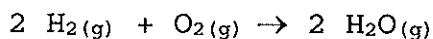
h)  $2 \text{SO}_{2(g)} + \text{O}_{2(g)} \rightarrow 2 \text{SO}_{3(g)}$

$$\Delta S < 0$$

i)  $\text{CaCl}_{2(s)} \rightarrow \text{Ca}^{2+}_{(aq)} + 2 \text{Cl}^{-}_{(aq)}$

$$\Delta S > 0$$

2. Use  $\Delta S_{\text{surr}}^\circ$  and  $\Delta S_{\text{Rxn}}^\circ$  to determine the  $\Delta S_{\text{univ}}^\circ$  at 298 K. Is the process spontaneous?



	$\Delta H_f^\circ$ (kJ/mol)	$S^\circ$ (J/mol.K)
$\text{H}_2$	0	+130.58
$\text{O}_2$	0	+205.0
$\text{H}_2\text{O}$	-241.82	+188.83

$$\Delta S_{\text{R}}^\circ = \sum n S_{\text{prod}}^\circ - \sum m S_{\text{react}}^\circ$$

$$= (2 \text{ mol} \times 188.83 \frac{\text{J}}{\text{mol} \cdot \text{K}}) - [(2 \text{ mol} \times 130.58 \frac{\text{J}}{\text{mol} \cdot \text{K}}) + (1 \text{ mol} \times 205.0 \frac{\text{J}}{\text{mol} \cdot \text{K}})]$$

$$\Delta S_{\text{R}}^\circ = -88.5 \text{ J/K}$$

$$\Delta H_{\text{R}}^\circ = \sum n \Delta H_{\text{f, prod}}^\circ - \sum m \Delta H_{\text{f, react}}^\circ$$

$$= (2 \text{ mol} \times -241.82 \frac{\text{kJ}}{\text{mol}}) - [(2 \text{ mol} \times 0 \frac{\text{kJ}}{\text{mol}}) + (1 \text{ mol} \times 0 \frac{\text{kJ}}{\text{mol}})]$$

$$\Delta H_{\text{R}}^\circ = -483.64 \text{ kJ} = -483,640 \text{ J}$$

$$\Delta S_{\text{surr}}^\circ = \frac{-\Delta H_{\text{R}}^\circ}{T} = \frac{-483,640 \text{ J}}{298 \text{ K}} = 1,620 \text{ J/K}$$

$$\Delta S_{\text{univ}}^\circ = \Delta S_{\text{sys}}^\circ + \Delta S_{\text{surr}}^\circ = -88.5 \text{ J/K} + 1,620 \text{ J/K} = \underline{\underline{1,530 \text{ J/K}}}$$